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Q1. During oxidation, a given substance loses electrons. Traditionally this loss was attributed to the presence of oxygen, or compounds containing oxygen. However, any substance that has the ability to gain electrons can cause oxidation to occur.

Q2. The original definition of reduction came from the reduction in mass of a metallic ore as it was transformed into a pure metal. The theoretical definition of reduction involves a substance gaining electrons. The key difference is the focus on a theoretical change in electron distribution as opposed to an observed change in mass. However, since the pure metal has a lower molar mass than the metallic oxide, the breakdown of the metal oxide is described by both definitions.

Q3. The oxidizing agent is the $Fe^{2+}(aq)$ ion, while the reducing agent is the Al(s).

Q4. A comparison of ionic charges among the reactants and products reveals that there has been no change in the distribution of electrons, meaning no redox reaction has taken place. Typically, in double replacement reactions, where all substances remain in the same ionic form, a redox reaction does not occur.

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Q5. **(a)** nonspontaneous

(b) spontaneous

(c) nonspontaneous

(d) spontaneous

Q6. Use **(b)** because the reaction will proceed spontaneously.

Q7. Any of the following metals will work: Fe, Zn, Cr, Al, Mg, Na, Ca, Ba, Li.

The reactions for Fe, Zn, Mg, Ca, and Ba, are similar:

$$Fe(s) + Cd^{2+}(aq) + 2Cl^{-}(aq) \rightarrow Fe^{2+}(aq) + 2Cl^{-}(aq) + Cd(s)$$

The reactions for Al and Cr are similar:

$$2Al(s) + 3Cd^{2+}(aq) + 2Cl^{-}(aq) \rightarrow 2Al^{3+}(aq) + 3Cl^{-}(aq) + 3Cd(s)$$

The reactions for Na and Li are also similar:

$$2Na(s) + Cd^{2+}(aq) + 2Cl^{-}(aq) \rightarrow 2Na^{+}(aq) + Cl^{-}(aq) + Cd(s)$$

Q8. The soluble salts could include chlorides of any metal ion listed below tin in the reactivity series. Below are three examples:

$$2Ag^{+}(aq) + Cl^{-}(aq) + Sn(s) \rightarrow 2Ag(s) + Sn^{2+}(aq) + 2Cl^{-}(aq)$$

$$Cu^{2+}(aq) + 2Cl^{-}(aq) + Sn(s) \rightarrow Cu(s) + Sn^{2+}(aq) + 2Cl^{-}(aq)$$

$$Pb^{2+}(aq) + 2Cl^{-}(aq) + Sn(s) \rightarrow Pb(s) + Sn^{2+}(aq) + 2Cl^{-}(aq)$$

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1. In an oxidation half-reaction, electrons appear on the product side (the right side, by convention). In an oxidation half-reaction, the reducing agent is the reactant, and the reducing agent gives up electrons. In a reduction half-reaction, the electrons appear on the reactant side (the left side, by convention). In a reduction half-reaction, the oxidizing agent is the reactant. Since the oxidizing agent gains electrons, electrons must also appear on the reactant side.

2. In a redox reaction, a molecule, atom, or ion gains electrons, while another molecule, atom, or ion loses electrons. Oxidation means losing electrons. For one chemical species to be oxidized, the other chemical species must accept electrons. Therefore, the chemical species that accepts electrons, facilitating oxidation, is the oxidizing agent.

3. Calcium ions will spontaneously react with Ba(s), K(s), Li(s).

$Ca^{2+}(aq) + Ba(s) \rightarrow Ba^{2+}(aq) + Ca(s)$

$Ca^{2+}(aq) + K(s) \rightarrow K^{+}(aq) + Ca(s)$

$Ca^{2+}(aq) + Li(s) \rightarrow Li^{+}(aq) + Ca(s)$

4. A combination reaction is any reaction in which two reactants combine to form a single product. Lithium is the most reactive of metals. Thus, it has the greatest tendency of all metals to give up its electrons in a chemical reaction. Therefore, if lithium reacts with another element or compound, it will be oxidized, and thus act as a reducing agent.

5. (a) The reaction will occur.

The complete balanced equation is:

$2AgNO_3(aq) + Cd(s) \rightarrow 2Ag(s) + Cd(NO_3)_2(aq)$

The ionic equation is:

$2Ag + (aq) + Cd(s) \rightarrow 2Ag(s) + Cd^{2+} + (aq)$

The net ionic equation is:

$2Ag^{+}(aq) + 2NO_3^{-}(aq) + Cd(s) \rightarrow 2Ag(s) + Cd^{2+}(aq) + 2NO_3^{-}(aq)$

(b) A reaction will not occur.

(c) The reaction will occur. The complete balanced reaction is:

$2Al(s) + 3HgCl_2(aq) \rightarrow 2AlCl_3(aq) + 3Hg(s)$

The net ionic equation is:

$2Al(s) + 3Hg^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Hg(s)$

The net ionic equation is:

$2Al(s) + 3Hg^{2+}(aq) + 6Cl^{-}(aq) \rightarrow 2Al^{3+}(aq) + 3Hg(s) + 6Cl^{-}(aq)$

6. Possible answers for each part are shown below. In general, equations must be balanced, oxidizing agents must gain electrons, and reducing agents must lose electrons. Oxidizing agents react with metals higher in the activity series; reducing agents react with metals lower in the activity series.

(a) $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

(b) $2Al(s) + 3Fe^{2+}(aq) \rightarrow 2Al^{3+}(aq) + 3Fe(s)$

(c) $Al(s) + Au^{3+}+(aq) \rightarrow Al^{3+}(aq) + Au(s)$

(d) $3Cu(s) + 2Au^{3+}+(aq) \rightarrow 3Cu^{2+}(aq) + 2Au(s)$

7. (a) $Na(\ell) + K^{+}(\ell) \rightarrow K(\ell) + Na^{+}(\ell)$

(b) The potassium ions are the oxidizing agents. Sodium is the reducing agent.

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Q9. The balanced half-reactions are as follows:

Oxidation: $Al(s) \rightarrow Al^{3+}(aq) + 3e^{-}$

Reduction: $Fe^{3+}(aq) + 3e^{-} \rightarrow Fe(s)$

Q10. The balanced half-reactions are as follows:

(a) Oxidation: $Fe(s) \rightarrow Fe^{2+}(aq) + 2e^{-}$

Reduction: $Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$

(b) Oxidation: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^{-}$

Reduction: $Ag^{+}(aq) + e^{-} \rightarrow Ag(s)$

Q11. The balanced half-reactions are as follows:

(a) Oxidation: $Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

Reduction: $Pb^{2+}(aq) + 2e^- \rightarrow Pb(s)$

(b) Oxidation: $Ag(s) \rightarrow Ag^+(aq) + e^-$

Reduction: $Au^{3+}(aq) + 3e^- \rightarrow Au(s)$

(c) Oxidation: $Zn(s) \rightarrow Zn^{2+} + 2e^-$

Reduction: $Fe^{3+}(aq) + 3e^- \rightarrow Fe(s)$

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1. (a) $5Ag(s) + MnO^{4-}(aq) + 8H^+(aq) \rightarrow 5Ag^+(aq) + Mn^{2+}(aq) + 4H_2O(\ell)$

(b) $Hg(\ell) + 4Cl^-(aq) + 2NO_3^-(aq) + 4H^+(aq) \rightarrow HgCl_4^{2-}N(s) + 2NO_2(g) + 2H_2O(\ell)$

(c) $AsH_3(g) + 4H_20(\ell) + 4Zn^{2+}(aq) \rightarrow H_3AsO_4(aq) + 8H^+(aq) + 4Zn(s)$

(d) $I_2(s) + H_2O(\ell) + 5OCl^-(aq) \rightarrow 2IO_3^-(aq) + 2H^+(aq) + 5Cl^-(aq)$

2. (a) $I^-(aq) + 6OH^-(aq) + 6MnO_4_(aq) \rightarrow 6MnO_4^{2-}(aq) + IO^{3-}(aq) + 3H_2O(\ell)$

(b) $3H_2O_2(aq) + 2OH^-(aq) + 2ClO_2(aq) \rightarrow 2ClO^- + 3O_2(g) + 4H_2O(\ell)$

(c) $2CrO_2^-(aq) + 6ClO^-(aq) + 2H_2O(\ell) \rightarrow 3Cl_2(g) + 4OH^-(aq) + 2CrO_4^{2-}(aq)$

(d) $4Al(s) + OH^-(aq) + 3NO^-(Aq) + 4H_2O(\ell) \rightarrow 3NH_3(g) + 4AlO_2^-(aq)$

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3. $2PbSO_4(aq) + 2H_2O(\ell) \rightarrow Pb(s) + PbO_2(aq) + 2H_2SO_4(aq)$

4. $3NO_2(g) + H_2O(\ell) \rightarrow NO(g) + 2HNO_3(aq)$

5. $Cl_2(g) + 2OH^-(aq) \rightarrow Cl^-(aq) + ClO^-(aq) + H_2O(\ell)$

6. $3I_3^-(aq) + 3H_2O(\ell) \rightarrow IO_3^-(aq) + 6H^+(aq) + 8I^-(aq)$

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1. (a) $Cr_2O_7^{2-} + 14H^+(aq) + 6e^- \rightarrow 2Cr^{3+}(aq) + 7H_2O(\ell)$; oxidation half-reaction

(b) $2S_2O_3^{2-}(aq) \rightarrow S_4O_6^{2-}(aq) + 2e^-$; oxidation half-reaction

(c) Arsenic is reduced, so this is a reduction half-reaction. First balance the atoms other than oxygen and hydrogen.

$4AsO_4(aq) \rightarrow As_4O_6(aq)$

Add water to balance the oxygen atoms.

$4AsO_4^{3-}(aq) \rightarrow As_4O_6(aq) + 10H_2O(\ell)$

Add hydrogen ions (acidic solution) to balance the hydrogen atoms.

$4AsO_4^{3-}(aq) + 20H^+(aq) \rightarrow As_4O_6(aq) + 10H_2O(\ell)$

Add electrons to balance the charges on both sides.

$AsO_4^{3-}(aq) + 20H^+(aq) + 8e^- \rightarrow As_4O_6(aq) + 10H_2O(\ell)$

(d) Bromine is oxidized, so this is an oxidation half-reaction. First balance the atoms other than oxygen and hydrogen.

$Br_2(\ell) \rightarrow 2BrO_3^-(aq)$

Add water to balance the oxygen atoms.

$Br_2(\ell) + 6H_2O(\ell) \rightarrow 2BrO_3^-(aq)$

Add hydrogen ions (acidic solution) to balance the hydrogen atoms.

$Br_2(\ell) + 6H_2O(\ell) \rightarrow 2BrO_3^-(aq) + 12H^+(aq)$

Adjust for basic conditions by adding hydroxide ions to both sides.

$$Br_2(\ell) + 6H_2O(\ell) + 12OH^-(aq) \rightarrow 2BrO_3^-(aq) + 12H^+(aq) + 12OH^-(aq)$$

Combine hydrogen and hydroxide ions and cancel any water found on both sides of the equation.

$Br_2(\ell) + 12OH^-(aq) \rightarrow 2BrO_3^-(aq) + 6H_2O(\ell)$

Add electrons to balance the charges on both sides.

$$Br_2(\ell) + 12OH^-(aq) \rightarrow 2BrO_3^-(aq) + 6H_2O(\ell) + 10e^-$$

2. (a) Balance this equation by balancing charge:

$3Co^{2+}(aq) + 2\,Au(s) \rightarrow 3Co(s) + 2Au^{3+}(aq)$

(b) Under acidic conditions, the balanced half-reactions are:

$Cu(s) \rightarrow Cu^{2+}(aq) + 2e^-$

$NO_3^-(aq) + 3e^- + 4H^+(aq) \rightarrow NO(g) + 2H_2O(\ell)$

Write half-reactions with the LCM.

$3Cu(s) \rightarrow 3Cu^{2+}(aq) + 6e^-$

$$2NO_3^-(aq) + 6e^- + 8H^+(aq) \rightarrow 2NO(g) + 4H_2O(\ell)$$

Add half-reactions, remove electrons and chemical species present on both sides:

$$3Cu(s) + 2NO_3^-(aq) + 8H^+(aq) \rightarrow 3Cu^{2+}(aq) + 2NO(g) + 4H_2O(\ell)$$

(c) Under basic conditions, the balanced half-reactions are:

$$Al(s) + 4OH^-(aq) \rightarrow AlO_2^-(aq) + 2H_2O(\ell) + 3e^-$$

$$NO_3^-(aq) + 6H_2O(\ell) + 8e^- \rightarrow NH_3(g) + 9OH^-(aq)$$

Write half-reactions with the LCM.

$$8Al(s) + 32OH^-(aq) \rightarrow 8AlO_2^-(aq) + 24e^- + 16H_2O(\ell)$$

$$3NO_3^-(aq) + 24e^- + 18H_2O(\ell) \rightarrow 3NH_3(g) + 27OH^-(aq)$$

Add half-reactions, remove electrons and chemical species present on both sides:

$$8Al(s) + 5OH^-(aq) + 3NO_3^-(aq) + 2H_2O(\ell) \rightarrow 8AlO_2^-(aq) + 3NH_3(g)$$

3. $Hg^+(aq) \rightarrow Hg^{2+}(aq) + e^-$

$Hg^+(aq) + e^- \rightarrow Hg(\ell)$

Combined, the net ionic equation would be:

$2Hg^+(aq) \rightarrow Hg^{2+}(aq) + Hg(\ell)$

4. (a) The balanced chemical reaction is:

$$3Fe_2O_3(s) + CO(g) \rightarrow 2Fe_3O_4(s) + CO_2(g)$$

(b) The balanced chemical reaction is:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

(c) The balanced chemical reaction is:

$$Fe_3O_4(s) + CO(g) \rightarrow 3FeO(s) + CO_2(g)$$

(d) The balanced chemical reaction is:

$$FeO(s) + CO(g) \rightarrow Fe(\ell) + CO_2(g)$$

5. Since electrons do not exist on their own, a gain or loss of electrons by one substance needs to correspond to a loss or gain of electrons by another substance. Thus, if one substance undergoes oxidation, another must simultaneously undergo reduction.

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7. (a) +3

(b) 0

(c) +6

(d) +5

(e) 0

(f) +2

8. (a) H = +1, S = +4, O = –2

(b) H = +1, O = –2

(c) H = +1, P = +5, O = –2

9. (a) O = +2

(b) O = –1

10. (a) Al = +3, H = +1, C = +4, O = –2

(b) N = –3, H = +1, P = +5, O = –2

(c) K = +1, H = +1, I = +7, O = –2

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11. (a) This is a redox reaction.

(b) This is not a redox reaction.

(c) This is a redox reaction.

(d) This is a redox reaction and a disproportionation reaction.

12. (a) Oxidizing agent: $H_2O_2(aq)$. Reducing agent: $Fe(OH)_2(s)$

(b) There is no oxidizing or reducing agent

(c) Oxidizing agent: $O_2(g)$. Reducing agent: $C_2H_6(g)$.

(d) $NO_2(g)$ is both the oxidizing and reducing agent.

13. The oxidation number of Br decreases from 0 in $Br_2(\ell)$ to –1 in $2Br^-(aq)$. This is reduction. $Br_2(\ell)$ is an oxidizing agent. The oxidation number of Cl increases from +3 in $ClO_2^-(aq)$ to +4 in $ClO_2(aq)$. This is oxidation. $ClO_2^-(aq)$ is a reducing agent.

14. The oxidation number of nickel decreases from +2 in the sulfide to 0 in its elemental form. This is a reduction. A reducing agent must be used to achieve this change. This is a redox reaction. The same analysis can be used with CuS and Cu.

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15. $CS_2(g) + 3O_2(g) \rightarrow CO_2(g) + 2SO_2(g)$

16. (a) $B_2O_3(aq) + 6Mg(s) \rightarrow 3MgO(s) + Mg_3B_2(aq)$

(b) $8H_2S(g) + 8H_2O(aq) \rightarrow S_8(s) + 16H_2O(\ell)$

17. (a) $Cr_2O_7^{2-}(aq) + 6Fe^-(aq) + 14H(aq) \rightarrow 2Cr^{3+}(aq) + 6Fe^{3+}(aq) + 7H_2O(\ell)$

(b) $I_2(g) + 10NO_3^-(aq) + 8H^+(aq) \rightarrow 2IO_3^-(aq) + 10NO_2(g) + 4H_2O(\ell)$

(c) $2PbSO_4(aq) + 2H_2O(\ell) \rightarrow Pb(s) + PbO_2(aq) + 2SO_4^{2-}(aq) + 4H^+(aq)$

18. (a) $3Cl^-(aq) + 2CrO_4^{2-}(aq) + H_2O(\ell) \rightarrow 3ClO^-(aq) + 2CrO_2^-(aq) + 2OH^-(aq)$

(b) $3Ni(s) + 2MnO_4^-(aq) + H_2O(\ell) \rightarrow 3NiO(s) + 2MnO_2(s) + 2OH^-(aq)$

(c) $I^-(aq) + 6Ce^{4+}(aq) + 6OH^-(aq) \rightarrow IO^{3-}(aq) + 6Ce^{3+}(aq) + 3H_2O(\ell)$

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1. (a) The reaction is a redox reaction because the oxidation numbers of the hydrogen and iodine atoms change.

$H_2(g) + I_2(g) \rightarrow 2HI(aq)$

0 0 +1 –1

(b) The reaction is not a redox reaction because the oxidation numbers of the atoms do not change.

$2NaHCO_3(aq) \rightarrow Na_2CO_3(aq) + H_2O(\ell) + CO_2(g)$

+1 +1 +4 –2 + +1 +4 –2 +1 –2 + +4 –2

(c) The reaction is not a redox reaction because the oxidation numbers of the atoms do not change.

$2HBr(aq) + Ca(OH)_2(aq) \rightarrow CaBr_2(aq) + 2H_2O(\ell)$

+1 –1 +2 –2 +1 +2 –1 +1–2

(d) The reaction is a redox reaction because the oxidation numbers of the phosphorus atom and two of the chlorine atoms change.

$PCl_5 \rightarrow PCl_3 + Cl_2(g)$

+5 – 1 +3 –1 0

2. Students may suggest definitions similar to the following:

- A redox reaction is a chemical reaction in which there is a change in the oxidation number of some or all of the atoms of the reactants.
- A redox reaction is a chemical reaction in which atoms of one element are oxidized and atoms of another element are reduced.

■ A redox reaction is a chemical reaction in which electrons are transferred among reactants. The number of electrons gained by the oxidizing agent must equal the number of electrons lost by the reducing agent.

3. Fluorine is the most electronegative element in the periodic table. Therefore, it does not "lose" electrons in chemical bonds. It has seven valence electrons, and thus forms bonds in which it gains one valence electron, satisfying the octet rule. Because it is in the second row of the periodic table, its outer occupied energy level cannot accommodate more than eight electrons (as bromine, iodine, or chlorine can). Therefore, it does not form bonds in which it gains more than one electron, and its oxidation number is always –1.

4. When one element combines with another element in a chemical reaction, the reaction is always a redox reaction. The atoms in uncombined elements have an oxidation number of 0. However, different elements have different electronegativities.. Therefore, when they combine in a compound, atoms of the less electronegative element are oxidized and atoms of the more electronegative element are reduced.

5. (a) Since the polyatomic ion contains oxygen, rule 4 applies. Therefore, students should assign an oxidation number of –2 to the oxygen atoms. Rule 7 also applies. Letting the oxidation number of sulfur be x, students can write:

$2x + 3(-2) = -2$

$x = +2$

Therefore, the oxidation number of sulfur in the thiosulfate ion is +2.

(b) Using a Lewis structure, students may arrive at two different oxidation numbers for sulfur using the following reasoning. The central sulfur atom usually has 6 electrons in its valence shell. The three oxygen atoms, however, have a greater electronegativity than sulfur. Therefore, they can be considered to have taken these electrons, leaving the central sulfur with a +6 oxidation number. The sulfur atom bonded to the central sulfur atom may be considered to share its electrons with the central sulfur. Therefore, it will have an oxidation number of –1. Since the central sulfur (+6) can now be considered to have one shared electron it will have a total oxidation number of +6 + (–1) = +5.

(c) Using the first method, the oxidation number for both sulfur atoms is +2. Using the second method, the oxidation number for one sulfur atom is –1, and for the other is +5. (Notice the average of the second result gives the first result).

(d) Using Lewis structures allows for the fact that atoms of the same element in a molecule or polyatomic ion may have different oxidation numbers, depending on the bonding arrangement. The disadvantage to this method is that it is more difficult and time consuming than using the rules.

(e) The advantage to using the oxidation number rules is that they apply to most cases and are quick and easy to use. The disadvantage to using these rules is that they do not allow for the possibility of different oxidation numbers for the same atoms in a molecule or polyatomic ion.

6. (a) $N_2 + 3H_2(g) \rightarrow 2NH_3(g)$

0 0 –3 +1

Nitrogen is the oxidizing agent, and hydrogen is the reducing agent.

(b) $NH_3(aq) + HNO_3 \rightarrow NH_4NO_3(aq)$

$-3+1 \quad +1\ +5\ -2\ -3\ +1+5\ -2$

The oxidation numbers of the atoms do not change. Therefore, the reaction is not a redox reaction.

7. (a) Because the reaction takes place in neutral conditions and there are no charges to worry about, the equation may be balanced by inspection. (Students should already be accustomed to balancing combustion equations by inspection.) Students should balance carbon first, then hydrogen, and finally oxygen. Students may also balance the equation using the oxidation number method.

$CH_3COOH(aq) + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(\ell)$

(b) Students may use either the half-reaction method or the oxidation number method to balance the equation. The oxidation number method is shown here.

$O_2(g) + H_2SO_3(aq) \rightarrow HSO_4^-(aq)$

$0 + \quad 1\ +4\ -2 \quad +1\ +6\ -2$

Sulfur's oxidation number increases by 2. Oxygen's oxidation number decreases by 2. Therefore, the ratio of sulfur atoms being oxidized to oxygen atoms being reduced must be 1:1.

$O_2(g) + 2H_2SO_3(aq) \rightarrow 2HSO_4^-(aq)$

Add hydrogen ions to balance atoms and charges. The resulting equation is already simplified.

$O_2(g) + 2H_2SO_3(aq) + \rightarrow 2HSO_4^-(aq) + 2H^+(aq) +$

8. (a) $NH_3(g) + Cl_2(g) \rightarrow NH_4Cl(aq) + N_2(g)$

$-3\ +1 \quad 0 \quad -3\ +1\ -1 \quad 0$

Nitrogen's oxidation number increases by 3. Chlorine's oxidation number decreases by 1.

Therefore, the ratio of nitrogen atoms being oxidized to chlorine atoms being reduced must be 1:3. Since 2 nitrogen atoms are oxidized, 6 chlorine atoms must be reduced.

$$NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(aq) + N_2(g)$$

Balance nitrogen atoms by inspection.

$$8NH_3(g) + 3Cl_2(g) \rightarrow 6NH_4Cl(aq) + N_2(g)$$

(b) $Mn_3O_4(aq) + Al(s) \rightarrow Al_2O_3 + Mn(s)$

$+8/3\ -2 \qquad 0 \qquad +3\ -2 \quad 0$

The oxidation number of manganese decreases by 8/3 and aluminium's oxidation number increases by 3. (The fraction occurs because the manganese oxide is considered to be composed of two different oxides, MnO and Mn_2O_3.) Aluminium is oxidized and each aluminium atom, on reduction, provides three electrons. The ratio of aluminium atoms oxidized to manganese atoms reduced must therefore be 8/3:3, or 8:9.

$$3Mn_3O_4(aq) + 8Al(s) \rightarrow Al_2O_3(aq) + Mn(s)$$

Balance the atoms by inspection.

$$3Mn_3O_4(aq) + 8Al(s) \rightarrow 4Al_2O_3(aq) + 9Mn(s)$$

9. The total increase in oxidation numbers corresponds to the number of electrons lost by oxidation. The total decrease in oxidation numbers corresponds to the number of electrons gained by reduction. In a reaction, net gain of electrons must be equal to net loss of electrons. Thus, the total increase in oxidation numbers must equal the total decrease in oxidation numbers.

10. The individual half-reactions are:

$$2H_2O(\ell) + NH_3(g) \rightarrow NO_2(g) + 7e^- + 7H^+(aq)$$

$$4H^+(aq) + 2O_2(g) + 4e^- \rightarrow 2H_2O(\ell)$$

The balanced half-reactions are:

$$8H_2O(\ell) + 4NH_3(g) \rightarrow 4NO_2(g) + 28e^- + 28H^+(aq)$$

$$28H^+(aq) + 14O_2(g) + 28e^- \rightarrow 14H_2O(\ell)$$

Giving:

$$4NH_3(g) + 14O_2(g) \rightarrow 4NO_2(g) + 6H_2O(\ell)$$

Which reduces to:

$$2NH_3(g) + 7O_2(g) \rightarrow 2NO_2(g) + 3H_2O(\ell)$$

The key assumption that students have to make is that an acidic environment is created in the ammonia half-reaction, whereas the oxygen reaction starts in an acidic environment. Since the hydrogen ions and extra waters cancel out, these assumptions do not affect the overall reaction.

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1. The permanganate ion is a good choice of oxidizing agent because it is very reactive, and it undergoes a colour change. The $MnO_4^-(aq)$ ion is dark purple and turns colourless as it is reduced to Mn^{2+} ions.
2. The starch is added to the mixture to act as a visual indicator.
3. If orange juice (or any antioxidant) has been exposed to the air, it becomes oxidized by oxygen. The concentration of ascorbic acid decreases as the compound reacts to form dehydroascorbic acid.
4. The chlorine oxidized the iodide ions back to triiodide ions, which regenerated some molecular iodine.
5. **(a)** Students may suggest that iodine could escape as a gas or react with oxygen or organic matter. Therefore, the concentration of the solution is unstable.

(b) Crush a tablet of vitamin C and dissolve it in water. Titrate with iodine. Record the endpoint. You now know the volume of iodine solution required to titrate the amount of vitamin C contained in the tablet, according to the package label.

(c) Over time, the vitamin C in the tablets is oxidized by oxygen in the air.

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Answers to Understanding Concepts Questions

1. (a) (i) $Zn(s) + 2AgNO_3(aq) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$

(ii) $3CoBr_2(aq) + 2Al(s) \rightarrow 3Co(s) + 2AlBr_3(aq)$

(iii) $Cd(s) + SnCl_2(aq) \rightarrow CdCl_2(aq) + Sn(s)$

(b) (i) Total ionic equation:

$$Zn(s) + 2Ag^+(aq) + 2NO_3^-(aq) \rightarrow Zn^{2+}(aq) + 2NO_3^-(aq) + 2Ag(s)$$

Net ionic equation:

$$Zn(s) + 2Ag^+(aq) \rightarrow Zn^{2+}(aq) + 2Ag(s)$$

(ii) Total ionic equation:

$$3Co^{2+}(aq) + 6Br^-(aq) + 2Al(s) \rightarrow 3Co(s) + 2Al^{3+}+(aq) + 6Br^-(aq)$$

Net ionic equation:

$$3Co^{2+}(aq) + 2A1(s) \rightarrow 3Co(s) + 2Al^{3+}(aq)$$

(iii) Total ionic equation:

$$Cd(s) + Sn^{2+}(aq) + 2Cl^-(aq) \rightarrow Cd^{2+}(aq) + 2Cl^-(aq) + Sn(s)$$

Net ionic equation:

$$Cd(s) + Sn^{2+}(aq) \rightarrow Cd^{2+}(aq) + Sn(s)$$

(c) (i) oxidizing agent: $Ag^+(aq)$; reducing agent: $Zn(s)$

(ii) oxidizing agent: $Co^{2+}(aq)$; reducing agent: $Al(s)$

(iii) oxidizing agent: $Sn^{2+}(aq)$; reducing agent: $Cd(s)$

(d) **(i)** oxidation half-reaction: $Zn(s) \rightarrow Zn^{2+}(aq) + 2e^-$

reduction half-reaction: $Ag^+(aq) + e^- \rightarrow Ag(s)$

(ii) oxidation half-reaction: $Al(s) \rightarrow Al^{3+}(aq) + 3e^-$

reduction half-reaction: $Co^{2+}(aq) + 2e^- \rightarrow Co(s)$

(iii) oxidation half-reaction: $Cd(s) \rightarrow Cd^{2+}(aq) + 2e^-$

reduction half-reaction: $Sn^{2+}(aq) + 2e^- \rightarrow Sn(s)$

2. **(a)** metallic element

(b) nonmetallic element

(c) nonmetallic element

(d) metallic element

3. **(a)** Ba +2; Cl –1

(b) Al +3; C –4

(c) K +1; C +2; N –3

(d) Li +1; N +3; O –2

(e) N –3; H +1; C +3; O –2

(f) S 0

(g) As +3; O –2

(h) V +5; O –2

(i) Xe +6; O –2; F –1

(j) S +2.5; O –2

4. In its ore, a metal is found in the oxidized form. In other words, it has a positive oxidation number. Pure metal, on the other hand, has an oxidation number of 0. Thus, when pure metal is extracted from ore, the oxidation number of the metal decreases—the metal is reduced.

5. **(a)** $O_2(g)$ is the oxidizing agent, and $C_6H_6(\ell)$ is the reducing agent.

(b) The reaction is not a redox reaction because the oxidation numbers do not change.

(c) $I_2(s)$ is the oxidizing agent, and $H_2(g)$ is the reducing agent.

(d) $KMnO_4(aq)$ is the oxidizing agent, and CuCl(s) is the reducing agent.

(e) Cu(s) undergoes oxidation, and $Ag^+(aq)+$ undergoes reduction.

(f) The reaction is not a redox reaction because the oxidation numbers do not change.

(g) $Mn^{2+}(aq)$ undergoes oxidation, and BiO_3^- undergoes reduction.

6. (a) oxidation numbers of vanadium are +5 in V_2O_5, VO_2^+, VO_3^-, VO_4^{3-}, $V_3O_9^{3-}$; and +4 in VO_2, VO^{2+}

(b) The reaction is not a redox reaction because the oxidation numbers do not change.

Answers to Applying Concepts Questions

7. The chlorite ion, ClO_2^-, is an example of an ion in which chlorine has an oxidation number of +3.

8. (a) Any synthesis reaction involving changes of oxidation numbers is a redox reaction. The easiest ones to come up with are ones that include elements in the reactants.

$$2H_2(g) + O_2(g) \rightarrow 2H_2O(\ell)$$

(b) Any synthesis reaction that does not involve electron transfer is not a redox reaction.

$$CaO(s) + H_2O(\ell) \rightarrow Ca(OH)_2(aq)$$

(c) Any decomposition that involves electron transfer is a redox reaction.

$$2C_7H_5(NO_2)_3(s) \rightarrow 12CO(g) + 2C(s) + 5H_2(g) + 3N_2(g)$$

(d) Any decomposition that does not involve electron transfer is not a redox reaction.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$

(e) Double displacement reactions are never redox reactions.

(f) Any double displacement reaction will work.

$FeS(s) + 2HCl(aq) \rightarrow H_2S(g) + FeCl_2(aq)$

9. (a) $2P_4(s) + 12H_2O(\ell) \rightarrow 5PH_3(s) + 3H_3PO_4(aq)$

(b) Phosphorus is both reduced and oxidized. Phosphorus atoms, with oxidation number 0, are reduced in $PH_3(s)$, with oxidation number –3. They are oxidized in $H_3PO_4(aq)$, with oxidation number +5.

10. (a) $2Al(s) + Fe_2O_3(s) \rightarrow 2Fe(s) + Al_2O_3(s)$

(b) The oxidizing agent is $Fe_2O_3(S)$; the reducing agent is $Al(s)$.

11. Students should describe a procedure similar to Investigation 12.A, in which they use single displacement reactions to decide whether tin or nickel is the better reducing agent. They will need samples of each metal, and solutions containing ions of the metals. They can then compare observations of tin metal in a solution of nickel ions, and nickel metal in a solution of tin ions. The metal that is the better reducing agent will react in a solution containing ions of the other metal.

12. Oxygen is consumed and carbon dioxide is produced. Therefore, some of the oxygen taken in by the body is reduced, indicating a redox reaction must have taken place.

13. In order for sulfide ions to act as oxidizing agents, they need to gain electrons. Since the sulfide ion has an octet of electrons, it is unlikely that the ion will gain more electrons to act as an oxidizing agent.

Answers to Solving Problems Questions

14. (a) $MnO_2(s) + Cl^-(aq) \rightarrow Mn^{2+}(aq) + Cl_2(g)$ (acidic conditions)

$2Cl^-(aq) \rightarrow Cl_2(g)$

$MnO_2(s) \rightarrow Mn^{2+}(aq)$

Balance the oxygen by adding water molecules.

$2Cl^-(aq) \rightarrow Cl_2(g)$

$MnO_2(s) \rightarrow Mn^{2+}(aq) + 2H_2O(\ell)$

Balance the hydrogen atoms by adding hydrogen ions.

$2Cl^-(aq) \rightarrow Cl_2(g)$

$MnO_2(s) + 4H^+(aq) \rightarrow Mn^{2+}(aq) + 2H_2O(\ell)$

Balance charge by adding electrons.

$2Cl^-(aq) \rightarrow Cl_2(g) + 2e^-$

$MnO_2(s) + 4H^+(aq) + 2e^- \rightarrow Mn^{2+}(aq) + 2H_2O(\ell)$

LCM is in place. Therefore, all that remains is to add the half-reaction equations.

$2Cl^-(aq) + MnO_2(s) + 4H^+(aq) \rightarrow Cl_2(g) + Mn^{2+}(aq) + 2H_2O(\ell)$

(b) $NO(g) + Sn(s) \rightarrow NH_2OH(aq) + Sn^{2+}(aq)$ (acidic conditions)

$Sn(s) \rightarrow Sn^{2+}(aq) + NO(g) \rightarrow NH_2OH(aq)$

Balance the hydrogen atoms by adding hydrogen ions.

$Sn(s) \rightarrow Sn^{2+}(aq)$

$NO(g) + 3H^+(aq) \rightarrow NH_2OH(aq)$

Balance charge by adding electrons.

$Sn(s) \rightarrow Sn^{2+}(aq) + 2e^-$

$NO(g) + 3H^+(aq) + 3e^- \rightarrow NH_2OH(aq)$

Determine the LCM for the electrons and rewrite the half–reactions.

$3Sn(s) \rightarrow 3Sn^{2+}(aq) + 6e^-$

$2NO(g) + 6e^- + 6H^+(aq) + \rightarrow 2NH_2OH(aq)$

Add the half-reaction equations.

$2NO(g) + 3Sn(s) + 6H^+(aq) \rightarrow 2NH_2OH(aq) + 3Sn^{2+}(aq)$

(c) $Cd^{2+}(aq) + V^{2+}(aq) \rightarrow Cd(s) + VO_3^-(aq)$ (acidic conditions)

$V^{2+}(aq) \rightarrow VO_3^-(aq)$

$Cd^{2+}(aq) \rightarrow Cd(s)$

Balance the oxygen atoms by adding water molecules.

$V^{2+}(aq) + 3H_2O(\ell) \rightarrow VO_3^-(aq)$

$Cd^{2+}(aq) + \rightarrow Cd(s)$

Balance the hydrogen atoms by adding hydrogen ions.

$V^{2+}(aq) + 3H_2O(\ell) \rightarrow VO_3^- + (aq) + 6H^+(aq)$

$Cd^{2+}(aq) \rightarrow Cd(s)$

Balance charge by adding electrons.

$V^{2+}(aq) + 3H_2O(\ell) \rightarrow VO_3^-(aq) + 6H^+(aq) + 3e^-$

$Cd^{2+}(aq) + 2e^- \rightarrow Cd(s)$

Determine the LCM for the electrons and rewrite the half-reactions.

$2V^{2+}(aq) + 6H_2O(\ell) \rightarrow 2VO_3^-(aq) + 6e^- + 12H^+(aq)$

$3Cd^{2+}(aq) + 6e^- \rightarrow 3Cd(s)$

Add the half-reaction equations.

$3Cd^{2+}(aq) + 2V^{2+}(aq) + + 6H_2O(\ell) \rightarrow 3Cd(s) + 2VO_3^-(aq) + 12H^+(aq)$

(d) $Cr(s) \rightarrow Cr(OH)_4^-(aq) + H_2(g)$ (basic conditions)

$Cr(s) \rightarrow Cr(OH)_4^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Balance the oxygen atoms by adding water molecules.

$Cr(s) + 4H_2O(\ell) \rightarrow Cr(OH)_4^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Balance the hydrogen atoms by adding hydrogen ions.

$Cr(s) + 4H_2O(\ell) \rightarrow Cr(OH)_4^-(aq) + 4H^+(aq)$

$2H^+(aq) + \rightarrow H_2(g)$

Add hydroxide ions to neutralize hydrogen ions.

$Cr(s) + 4H_2O(\ell) + 4OH^-(aq) \rightarrow$
$Cr(OH)_4^-(aq) + 4H^+(aq) + 4OH^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Combine the hydrogen and hydroxide ions and remove any identical molecules that are common to both sides of the equation.

$Cr(s) + 4OH^-(aq) \rightarrow Cr(OH)_4^-(aq)$

$2H^+(aq) \rightarrow H_2(g)$

Balance charge by adding electrons.

$Cr(s) + 4OH^-(aq) \rightarrow Cr(OH)_4^-(aq) + 3e^-$

$2H^+(aq) + 2e^- \rightarrow H_2(g)$

Determine the LCM for the electrons and rewrite the half-reactions.

$2Cr(s) + 8OH^-(aq) \rightarrow 2Cr(OH)_4^-(aq) + 6e^-$

$6H^+(aq) + 6e^- \rightarrow 3H_2(g)$

Add the half-reaction equations. (Combine the hydrogen ions with the hydroxide ions to form water on the reactant side).

$2Cr(s) + 6H_2O(\ell) + 2OH^-(aq) \rightarrow$
$2Cr(OH)_4^-(aq) + 3H_2(g)$

(e) $S_2O_3^{2-}(aq) + NiO_2(s) \rightarrow Ni(OH)_2(s) + SO_3^{2-}(aq)$ (basic conditions)

Nickel is reduced (+4 to +2) and sulfur is oxidized (+2 to +4).

$S_2O_3^{2-}(aq) \rightarrow SO_3^{2-}(aq)$

$NiO_2(s) \rightarrow Ni(OH)_2(s)$

Balance the number of atoms other than hydrogen and oxygen.

$S_2O_3^{2-}(aq) \rightarrow 2SO_3^{2-}(aq)$

$NiO_2(s) \rightarrow Ni(OH)_2(s)$

Balance the number of oxygen atoms by adding water.

$3H_2O(\ell) + S_2O_3^{2-}(aq) \rightarrow 2SO_3^{2-}(aq)$

$NiO_2(s) \rightarrow Ni(OH)_2(s)$

Balance the number of hydrogen atoms by adding hydrogen ions.

$3H_2O(\ell) + S_2O_3^{2-}(aq) \rightarrow 2SO_3^{2-}(aq) + 6H^+(aq)$

$2H^+(aq) + NiO_2(s) \rightarrow Ni(OH)_2(s)$

Add hydroxide ions to neutralize hydrogen ions.

$$3H_2O(\ell) + S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 6H^+(aq) + 6OH^-(aq)$$

$$2OH^-(aq) + 2H^+(aq) + NiO_2(s) \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$$

Combine the hydrogen and hydroxide ions and remove water molecules common to both sides of the equation.

$S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 3H_2O(\ell)$

$2H_2O(\ell) + NiO_2(s) \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$

Balance charge by adding electrons.

$$S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 3H_2O(\ell) + 4e^-$$

$$2H_2O(\ell) + NiO_2(s) + 2e^- \rightarrow Ni(OH)_2(s) + 2OH^-(aq)$$

Determine the LCM for the electrons and rewrite the half-reactions.

$$S_2O_3^{2-}(aq) + 6OH^-(aq) \rightarrow 2SO_3^{2-}(aq) + 3H_2O(\ell) + 4e^-$$

$$4H_2O(\ell) + 2NiO_2(s) + 4e^- \rightarrow 2Ni(OH)_2(s) + 4OH^-(aq)$$

Add the half-reaction equations and remove water molecules and hydroxide ions common to both sides of the equation.

$$S_2O_3^{2-}(aq) + H_2O(\ell) + 2OH^-(aq) + 2NiO_2(s) \rightarrow 2SO_3^{2-} + 2Ni(OH)_2(s)$$

(f) $Sn^{2+}(aq) + O_2(g) \rightarrow Sn^{4+}(aq)$ (basic conditions)

$$Sn^{2+}(aq) \rightarrow Sn^{4+}(aq)$$

$$O_2(g) + 2H_2O(\ell) \rightarrow 4OH^-(aq)$$

The number of atoms, including hydrogen and oxygen, are balanced. Balance charge and determine the LCM for the electrons and rewrite the half-reactions.

$$2Sn^{2+}(aq) \rightarrow 2Sn^{4+}(aq) + 4e^-$$

$$O_2(g) + 2H_2O(\ell) + 4e^- \rightarrow 4OH^-(aq)$$

Add the half-reaction equations.

$$2Sn^{2+}(aq) + O_2(g) + 2H_2O(\ell) \rightarrow 2Sn^{4+}(aq) + 4OH^-(aq)$$

15. (a) $SiCl_4(aq) + Al(s) \rightarrow Si(s) + AlCl_3(aq)$

$+4\ -1 \qquad 0 \qquad 0 \qquad +3\ -1$

Silicon gains four electrons and aluminium loses three electrons. The ratio of silicon tetrachloride to aluminium is 3:4.

$$3SiCl_4(aq) + 4Al(s) \rightarrow 3Si(s) + 4AlCl_3(aq)$$

(b) $PH_3(aq) + O_2(g) \rightarrow P_4O_{10}(s) + H_2O(\ell)$

$-3+1 \quad 0 \quad +5\,-2 \quad +1\,-2$

The oxidation number of phosphorus increases by 8. The oxidation number of oxygen decreases by 2. Therefore, the ratio of oxygen atoms to phosphorus atoms must be 8:2, or 4:1. Since there are already 4 phosphorus atoms, the ratio is 16:4.

$4PH_3(s) + 8O_2(g) \rightarrow P_4O_{10}(s) + H_2O(\ell)$

To complete the equation, balance the oxygen atoms on both sides.

$4PH_3(s) + 8O_2(g) \rightarrow P_4O_{10}(s) + 6H_2O(\ell)$

(c) $I_2O_5(s) + CO(g) \rightarrow I_2(s) + CO_2(g)$

$+5\,-2 \quad +2\,-2 \quad 0 \quad +4-2$

The oxidation number of iodine decreases by 5. The oxidation number of carbon increases by 2.

Therefore, the ratio of carbon to iodine is 5:2.

$I_2O_5(s) + 5CO(g) \rightarrow I_2(s) + 5CO_2(g)$

(d) $SO_3^{2-}(aq) + O_2(g) \rightarrow SO_4^{2-}(aq)$

$+4\,-2 \quad 0 \quad +6\,-2$

The oxidation number of sulfur increases by 2. The oxidation number of oxygen decreases by 2. Therefore, the ratio of oxygen reduced to sulfur oxidized is 1:1.

$SO_3^{2-}(aq) + O_2(g) \rightarrow SO_4^{2-}(aq)$

Balance oxygen atoms.

$SO_3^{2-}(aq) + \frac{1}{2}O_2(g) \rightarrow SO_4^{2-}(aq)$, or

$2SO_3^{2-}(aq) + O_2(g) \rightarrow 2SO_4^{2-}(aq)$

16. (a) Nitrogen is reduced (+4 to +3) and nitrogen is oxidized (+4 to +5) under acidic conditions.

$NO_2(g) \rightarrow NO_3^-(aq)$; reduction reaction

$NO_2(g) \rightarrow NO_2^-(aq)$; oxidation reaction

Students should balance the equation via the half-reaction method, resulting in the following net ionic equation:

$$2NO_2(g) + H_2O(\ell) \rightarrow NO_2^-(aq) + NO_3^-(aq) + 2H^+(aq)$$

(b) Chlorine is oxidized (0 to +1) and it is also reduced (0 to –1) under basic conditions.

$Cl_2 \rightarrow 2Cl^-(aq)$; the reduction reaction

$Cl_2(g) \rightarrow ClO^-(aq)$; the oxidation reaction.

Students should balance the equation via the half-reaction method, resulting in the following net ionic equation:

$$Cl_2(g) + 2OH^-(aq) \rightarrow ClO^- + (aq) + H_2O(\ell) + Cl^-(aq)$$

17. (a) First, students should balance the equation by inspection.

$$2Co^{3+}(aq) + Cd(s) \rightarrow 2Co^{2+}(aq) + Cd^{2+}(aq)$$

In this case, students can just add the spectator ions, with the following result:

$$2Co(NO_3)_3(aq) + Cd(s) \rightarrow 2Co(NO_3)_2(aq) + Cd(NO_3)_2(aq)$$

(b) In this equation, silver is reduced and sulfur is oxidized under acidic conditions. Students should follow the half-reaction method to obtain the following equation:

$$2Ag^+(aq) + SO_2(g) + 2H_2O(\ell) \rightarrow 2Ag(s) + SO_4^{2-} + (aq) + 4H^+(aq)$$

Students can now include the spectator ions and states.

$$2AgNO_3(aq) + SO_2(g) + 2H_2O(\ell) \rightarrow 2Ag(s) + H_2SO_4(aq) + 2HNO_3(aq)$$

(c) Aluminium is oxidized (0 to +3) and chromium is reduced (+6 to +3) in a basic solution. Students should follow the half-reaction method to obtain the following equation:

$$Al(s) + CrO_4^{2-}(aq) + 4H_2O(\ell) \rightarrow Al(OH)_3(s) + Cr(OH)_3(s) + 2OH^-(aq)$$

Students can now add the spectator ions and states.

$$Al(s) + Na_2CrO_4(aq) + 4H_2O(\ell) \rightarrow Al(OH)_3(s) + Cr(OH)_3(s) + 2NaOH(aq)$$

18. (a) $I_2(s) + 10HNO_3(aq) \rightarrow 2HIO_3(aq) + 10NO_2(g) + 4H_2O(\ell)$

(b) $$\text{mass iodine} = \frac{28.0\ \text{L}\ NO_2(g)}{22\ 4\ \text{L/mol}\ NO_2(g)} \times \frac{1\ \text{mol}\ I_2(s)}{10\ \text{mol}\ NO_2(g)} \times \frac{253.8\ \text{g}\ I_2(s)}{1\ \text{mol}\ I_2(s)}$$

$$= 3.17\ \text{g}\ I_2(s)$$

19. (a) Students should use the half-reaction or oxidation number method to balance the reaction, remembering to use basic conditions. The solution is:

$$P_4(s) + 3H_2O(\ell) + 3OH^-(aq) \rightarrow 3H_2PO^{2-}(aq) + PH_3(g)$$

(b) Phosphorus is both oxidized (0 to +1) and reduced (0 to –3).

(c) $$10.0\ \text{kg}\ P_4(s) \times \frac{1000\ \text{g}}{1\ \text{kg}} \times \frac{\text{mol}\ P_4(s)}{123.88\ \text{g}} \times \frac{1\ \text{mol}\ PH_3(g)}{1\ \text{mol}\ P_4(s)} \times \frac{34\ \text{g}}{1\ \text{mol}\ PH_3(g)}$$

$$= 2744\ \text{g}\ PH_3(g) \text{ or } 2.74 \times 10^3\ \text{g}\ PH_3(g)$$

Answers to Making Connections Questions

20. (a) Na +1; Al +3; O –2; H +1; C +4

(b) $NaAl(OH)_2CO_3(s) + 4HCl(aq) \rightarrow NaCl(aq) + AlCl_3(aq) + 3H_2O(\ell) + CO_2(g)$

(c) The oxidation numbers were not particularly useful, since the reaction is not a redox reaction.

(d) The reaction is a neutralization reaction, not a redox reaction.

(e) Student answers will be based on the product they choose. Most antacids contain sodium, calcium, aluminium, or magnesium salts.

21. (a) $2KClO_3(s) \rightarrow 2KCl(s) + 3O_2(g)$

$S(s) + O_2(g) \rightarrow SO_2(g)$

(b) In the first reaction the oxidizing agent and the reducing agent is $KClO_3(s)$. In the second reaction oxygen is the oxidizing agent and sulfur is the reducing agent.

(c) No element in potassium chlorate undergoes disproportionation. Each element appears only once in the formulas of the reactants. Therefore, each element undergoes either reduction or oxidation, not both.

(d) The safety match was invented in 1844 by Swedish professor Gustaf Erik Pasch. Until then, matches had been made using poisonous yellow phosphorus. Their toxicity was not their only drawback. Because the

matches could light when struck on any surface, they had a tendency to burst into flames unexpectedly. For his safety match, Pasch used red phosphorus instead of yellow phosphorus. Instead of adding the phosphorus to the match, Pasch impregnated a sandpaper strip on the side of the matchbox with the phosphorus. The match, which had the other reactants on its head, would only light when struck on the special sandpaper strip.

22. (a) The nitrogen is oxidized (–3 to +5) and the oxygen is reduced (0 to –2). Students should assume an acid environment, and balance the equation by the method of their choice.

$$NH_4^+(aq) + 2O_2(g) \rightarrow NO_3^- + (aq) + H_2O(\ell) + 2H^+(aq)$$

(b) Student answers should include references to relevant chemistry, as well as economic and social considerations.